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ABSTRACT

This booklet is part of an American Astronomical Society curriculum project designed to provide teaching materials to teachers of secondary school chemistry, physics, and earth science. A Basic Topics section discusses atomic structure, emphasizing states of matter at high temperature and spectroscopic analysis of light from the stars. A section on Intermediate and Advanced Topics provides greater detail and more mathematical analysis of spectroscopy theory, including the quantum theory of electromagnetic radiation, electron orbits, and the classification of stellar spectra. Appendices include a glossary of unfamiliar terms, references and teaching aids, suggested class exercises, and questions and answers. (MH)

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Atoms and Astronomy

SE 021 773

NASA
National Aeronautics and Space Administration

ATOMS IN ASTRONOMY

A curriculum project of the American Astronomical Society,

prepared with the cooperation of the

National Aeronautics and Space Administration and the National Science Foundation

by

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National Aeronautics and Space Administration
Washington, DC 20546 September 1976

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PREFACE

In the past half century astronomers have provided mankind with a new view of the universe, with glimpses of the nature of infinity and eternity that beggar the imagination. Particularly, in the past decade, NASA's orbiting spacecraft as well as ground-based astronomy have brought to man's attention heavenly bodies, sources of energy, stellar and galactic phenomena, about the nature of which the world's scientists can only surmise.

Esoteric as these new discoveries may be, astronomers look to the anticipated Space Telescope to provide improved understanding of these phenomena as well as of the new secrets of the cosmos which they expect it to unveil. This instrument, which can observe objects up to 30 to 100 times fainter than those accessible to the most powerful Earth-based telescopes using similar techniques, will extend the use of various astronomical methods to much greater distances. It is not impossible that observations with this telescope will provide glimpses of some of the earliest galaxies which were formed, and there is a remoter possibility that it will tell us something about the edge of the universe.

The researches of the past 10 years, plus the possibility of even more fundamental discoveries in the next decade, are fascinating laymen and firing the imagination of youth. NASA's inquiries into public interest in the space program show that a major source of such interest is stellar and galactic astronomy. NASA's enabling Act, the Space Act of 1958, lists a primary purpose of NASA, "the expansion of human knowledge of phenomena in the atmosphere and space"; the Act requires of NASA that "it provide for the widest practicable and appropriate dissemination of information concerning its activities and the results of those activities."

In the light of the above, NASA is publishing for science teachers, particularly teachers of secondary school chemistry, physics, and Earth science, the following four booklets prepared by the American Astronomical Society (AAS) with the cooperation of NASA:

The Supernova, A Stellar Spectacle, by Dr. W. C. Straka, Department of Physics, Jackson State University, Jackson, Mississippi.

Extragalactic Astronomy, The Universe Beyond our Galaxy by Dr. Kenneth C. Jacobs, Department of Astronomy, University of Virginia, Charlottesville, Virginia.

Chemistry Between the Stars, by Dr. Richard H. Gammon, National Radio Astronomy Observatory, Charlottesville, Virginia.

Atoms in Astronomy, by Dr. Paul A. Blanchard, Theoretical Studies Group, NASA Goddard Space Flight Center, Greenbelt, Maryland.

The National Science Foundation has cooperated in this project by funding for the AAS a High School Astronomy Education Workshop in June 1974 at the University of Richmond in order to give the manuscripts a thorough pedagogic review in terms of curricular relevance and classroom use. The resulting publications provide exciting accounts of recent discoveries in the cosmos, and of the nature of the scientific thought and techniques by which scientists are trying to understand these discoveries.

NASA expresses its appreciation to the authors and to the members of the AAS Task Group on Education in Astronomy (TGEA), whose enthusiasm and energy carried the project to completion, particularly to Dr. Gerrit L. Verschuur, Director of the Fiske Planetarium, University of Colorado, who served as Director of the project; Dr. Donat G. Wentzel, Astronomy Program, University of Maryland, initiator of the project; Dr. Paul H. Knappenberger, Jr., Director, the Science Museum of Virginia and Chairman of the TGEA, who served as Workshop Director, and Herman M. Gurin, Executive Officer of the American Astronomical Society. To those who were enrolled in the Workshop and to others whose judgments and suggestions helped give the manuscripts the necessary scientific and curricular validity, NASA is grateful.

Appreciation is also expressed to the National Science Foundation for its support of the Workshop, to the University of Richmond for its cordial and efficient service as host to the Workshop, and to the NASA Goddard Space Flight Center for its assistance in making possible the publication of this project.

Technical Monitor for the project was Dr. Nancy W. Boggess, Staff Astronomer, Astrophysics Division, Office of Space Science, NASA. Coordinators of the project were Dr. Nancy G. Roman, Chief, Astronomy/Relativity, Office of Space Science, NASA, and Dr. Frederick B. Tuttle, Director, Educational Programs Division, Office of Public Affairs, NASA. Assisting until his retirement in June 1974, was Mr. Myrl H. Ahrendt, Instructional Materials Officer, Educational Programs Division.

September 1976

National Aeronautics and Space Administration
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FOREWORD

In many respects, astronomy may be regarded as a branch of physics. In no area is this connection more relevant or obvious than in the field of astronomical spectroscopy. Hence, more than half of this brochure is devoted to those topics in electromagnetic radiation and atomic physics needed for an understanding of astronomical applications. Spectroscopy is of great importance for astronomy since virtually all of our information about the Sun, the stars, and other bodies in the universe relies upon an analysis of the electromagnetic radiation we receive from them.

Although intended primarily for teachers, this brochure has been written so that it can be distributed to students if desired. The first section, Basic Topics, would be suitable for a ninth-grade general science class; the style is simple and repetitive, and no mathematics or physics background is required. The second section, Intermediate and Advanced Topics, requires a knowledge of the material in the first section and assumes a generally higher level of achievement and motivation on the part of the student. These latter topics might fit well into junior-level physics, chemistry, or earth-science courses.

Paul Blanchard

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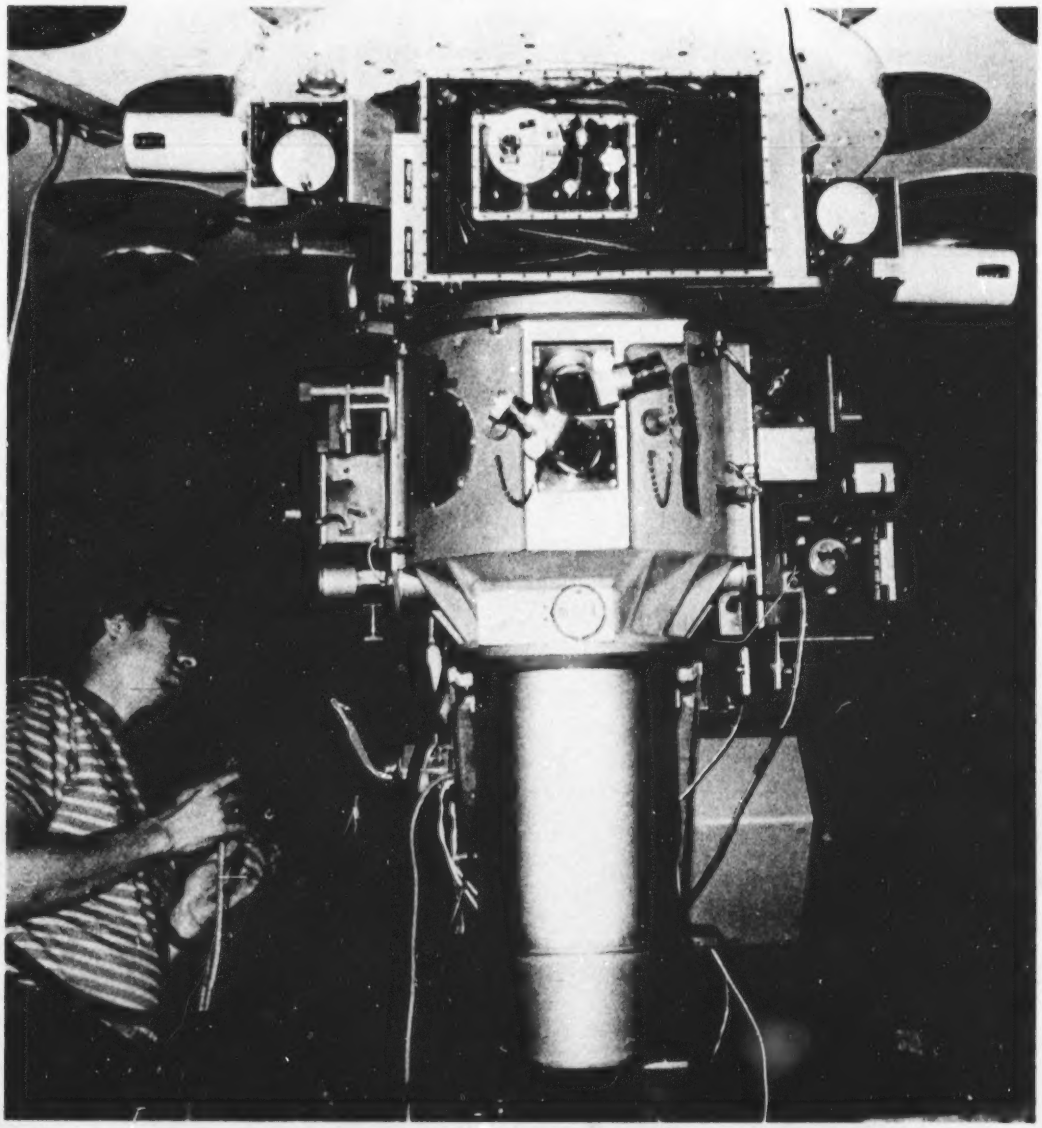
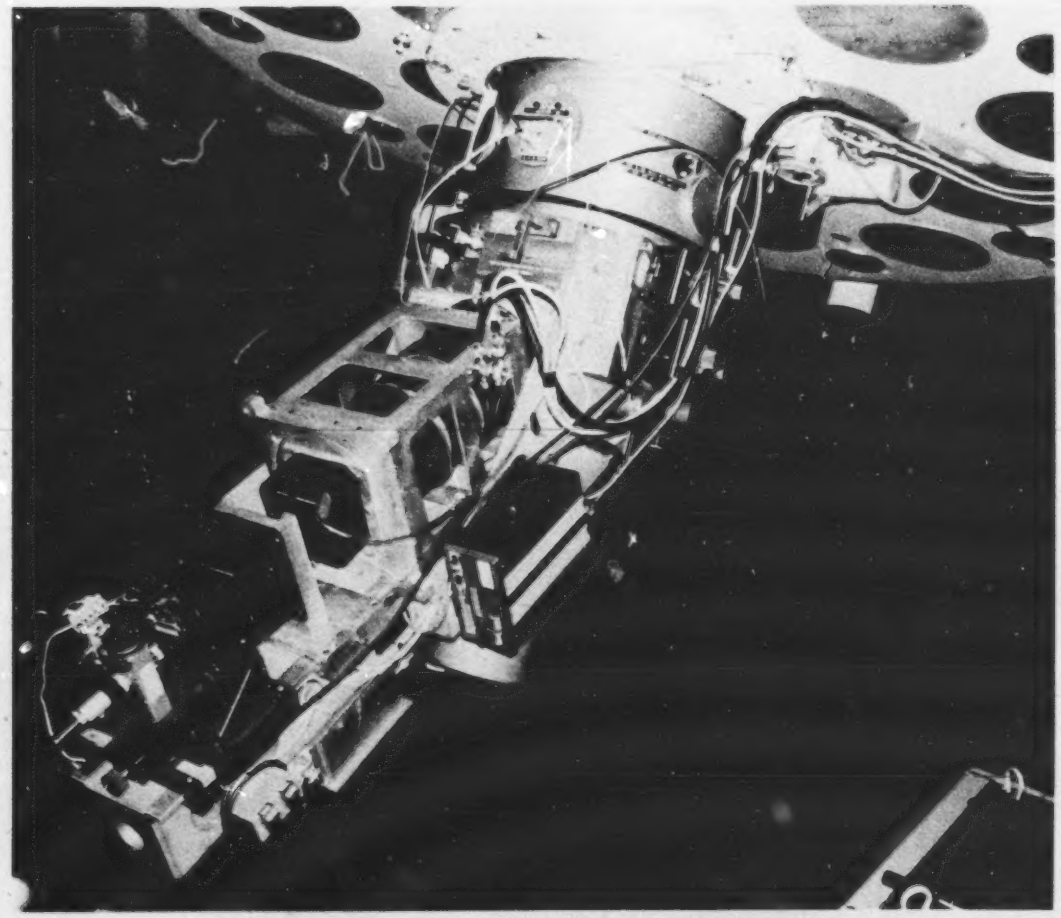


Astronomers at the Kitt Peak National Observatory prepare for observations with an image-tube spectrograph, used to record spectra of faint astronomical objects. This instrument is attached to the bottom of the tube containing the 2.1-m (84-in) telescope mirror.



The optical bench of the image-tube spectrograph shown in the preceding photograph (Kitt Peak National Observatory). The image tube, used to amplify weak light signals, protrudes from the spectrograph, followed by transfer lenses and photographic plate holder.

CHAPTER I

BASIC TOPICS

A. What are Atoms?

1. Atomic structure. As background, let us first recall the difference between a chemical element and a chemical compound. Chemical *elements* are those substances that cannot be broken down into simpler ones by chemical reactions, and so they are the simplest substances to display definite chemical properties. Some familiar elements are hydrogen (H), helium (He), carbon (C), nitrogen (N), oxygen (O), silver (Ag), and gold (Au). Chemical *compounds*, by contrast, are composed of the more basic chemical elements. For example, water (H_2O) is a compound, composed of hydrogen and oxygen.

An *atom* is the smallest unit of a chemical *element* that still possesses the chemical properties of that element. A *molecule* is the smallest unit of a chemical *compound* that still possesses the chemical properties of that compound. When a compound is created from elements in a chemical reaction, the atoms of the elements join together to form a molecule of the compound. Thus, molecules are simply atoms joined together.

An atom consists of a very small, dense *nucleus* surrounded by a number of rapidly moving *electrons*. Electrons are light, negatively charged particles that are familiar to us from the study of electricity; an electric current is just a flow of electrons, and it is these same particles that form the outer part of an atom. The nucleus, in the center of the atom, is itself composed of two kinds of particles – *protons* and *neutrons* – which are about 1800 times heavier than electrons. The protons carry a positive electric charge whereas the neutrons have no electric charge. Therefore, the nucleus as a whole is positively charged and there is an electrical force of attraction between the nucleus and the electrons that binds the atom together. The electrons do not fall into the nucleus, however, because they are in such rapid motion around it. Roughly speaking, we can say that the electrons are *in orbit* around the nucleus, although this analogy with planetary motion is not entirely correct.

As measured by the size of the cloud of orbiting electrons, atoms are extremely small. For example, it would take about 200 million hydrogen atoms, placed side by side, to stretch one centimeter. The nucleus is much smaller – about 100 000 times smaller than the cloud of electrons orbiting around it. Since the electrons themselves are smaller than the protons and neutrons that compose the nucleus, most of an atom is really empty space!

The atoms corresponding to different elements differ from each other in two ways: in the numbers of electrons they contain and in the numbers of protons and neutrons in the nucleus. Hydrogen, the simplest and lightest of all elements, has the simplest atom of all: a single electron in motion about a one-proton nucleus. The next simplest atom is that of helium, which has two electrons and a nucleus composed of two protons and two neutrons. The third simplest atom is that of lithium – a soft, silvery metal – which has three electrons and a nucleus of three protons and four neutrons (fig. 1). The most complicated, naturally occurring atom is that of uranium, which has 92 electrons and an equal number of protons in its nucleus.

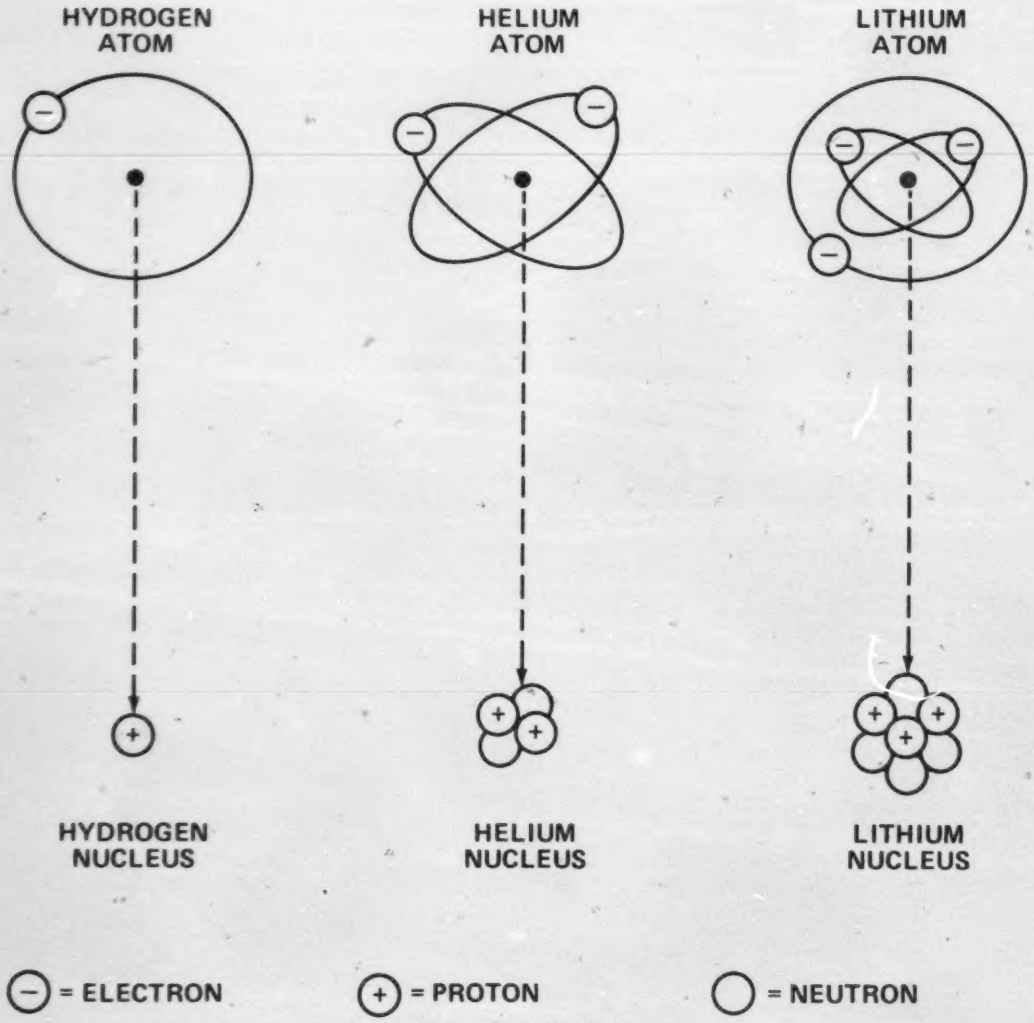


Figure 1 – Schematic diagrams of the three simplest atoms (hydrogen, helium, lithium) and their nuclei.

2. States of matter at high temperature. If a source of energy is available, such as that supplied by a chemical reaction, molecules can be broken down into their component atoms. Can atoms themselves be broken down into electrons and nuclei? The answer is "Yes," but this process also requires a source of energy. It is not very difficult to remove one or two electrons from an atom that already has a number of them. The removal of an electron from an atom is called *ionization*, and the energy liberated in many chemical reactions is sufficient to cause the ionization of an atom. However, removing most or all of the electrons from an atom requires a great amount of energy.

Under conditions of high temperature, the energy needed to ionize atoms is supplied by the motions of the atoms themselves. At high temperatures, all materials vaporize and become gases if they are not gases already. In a gas, the atoms are free to move around and are constantly colliding with each other. The higher the temperature, the faster the atoms move and the more violent are the collisions. At the surfaces of some stars – and in the extremely hot centers of all stars – these collisions are sufficiently violent to tear apart the atoms into electrons and nuclei. Under such conditions, matter can no longer be said to be composed of atoms; rather, it is composed of free electrons and atomic nuclei.

Finally, can the atomic nuclei themselves be broken down into protons and neutrons? The answer again is "Yes," but this process requires amounts of energy far greater than those required to break down atoms. The study of nuclei and their reactions forms a very special branch of physics called *nuclear physics.* At very high energies, nuclei can react with each other to form new nuclei or other types of nuclear particles. In the centers of stars, temperatures are high enough (millions of degrees) to permit such *nuclear reactions* to occur. We now know that these nuclear reactions are responsible for the energy output of the Sun and other stars.

B. Light and Spectroscopy

1. Newton's prism. An important property of atoms is that they can emit and absorb energy in the form of light. From a study of light emitted or absorbed by the atoms of a gas, scientists can learn much about the composition, temperature, and density of that gas. The analysis of light is called *spectroscopy.*

The great English scientist Isaac Newton (1642-1727) was the first to show that white light is actually a mixture of all the colors in the rainbow. Newton allowed a beam of sunlight to pass through a hole in a window shutter, then through a prism of glass, and finally fall upon a white screen beyond the prism (fig. 2). He found that the prism spreads out the white light in a row of colors called a *spectrum.* The red light is bent least by the prism, and the blue and violet lights are bent most. The colors appear in the order: violet, blue, green, yellow, orange, and red. Newton correctly concluded that all of these colors are present in white light and that the prism merely separates them.

Newton further showed that no color of the spectrum can be broken down into other colors. For example, if green light is passed through the prism, only green light comes out the other side; no other colors are observed. If, on the other hand, purple light – a mixture of red and blue – is used, these two colors will be separated by the prism. In general, the prism can be used to show what colors are present in any beam of light.

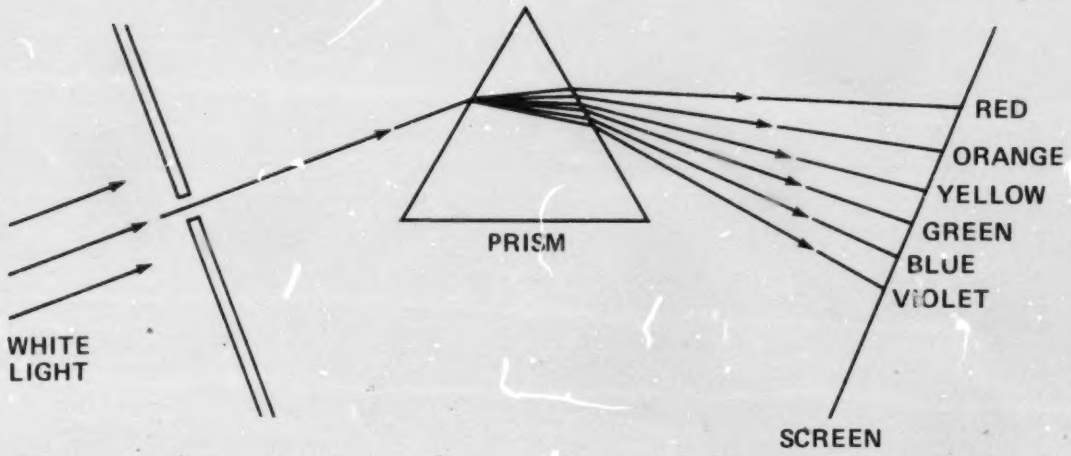


Figure 2 – Newton's prism experiment.

In experiments of this sort, it is convenient to use a *slit* to admit light to the prism, rather than the small hole which Newton first used. Figure 2 is drawn so that we are looking down the length of the slit, which is parallel to the axis (long dimension) of the prism. We then get a *band* of colors on the screen, rather than a narrow row. This band of colors has a height equal to the height of the slit. If a single color of light is passed through the slit, a *line* of that color appears on the screen behind the prism. Such a line of color is called a *spectral line.* For example, green light will produce a green spectral line in the green region of the spectrum. Purple light will produce two spectral lines – one red and one blue – at opposite ends of the spectrum.

We have seen that white light consists of a mixture of all the colors of the spectrum. White light is therefore said to have a *continuous spectrum,* because the colors run smoothly into each other and no colors are missing. Experimentation shows that any solid, hot, glowing body – or any highly compressed, hot, dense gas – gives off a continuous spectrum of light radiation. A good example is the light from an ordinary incandescent light bulb.

2. Emission lines. Some sources of light do not, however, have a continuous spectrum. Hot, glowing, *thin gases* are the most important examples of such sources. (By a "thin" gas we mean one of low density, one not highly compressed.) If the light from a hot, thin gas is passed through a slit and then a prism, one observes on the screen a pattern of bright, individual spectral lines. This kind of spectrum is called a *bright-line spectrum* or *emission-line spectrum.* Obviously, only certain colors of light are called emitted – those colors that appear in the spectrum as bright emission lines. The particular pattern of spectral lines depends upon the gas. Experiments first performed in the 19th century showed that *no two different gases have exactly the same pattern of spectral lines.* The emission-line spectrum of a hot, thin gas is therefore like the signature of a person – no two are exactly alike. The spectral emission lines of a gas may thus be used to identify that gas wherever it is present, whether in the laboratory or in astronomical sources of light radiation. Figure 3

illustrates the difference between the patterns of emission lines in the spectra of hydrogen (H) gas and mercury (Hg) vapor.

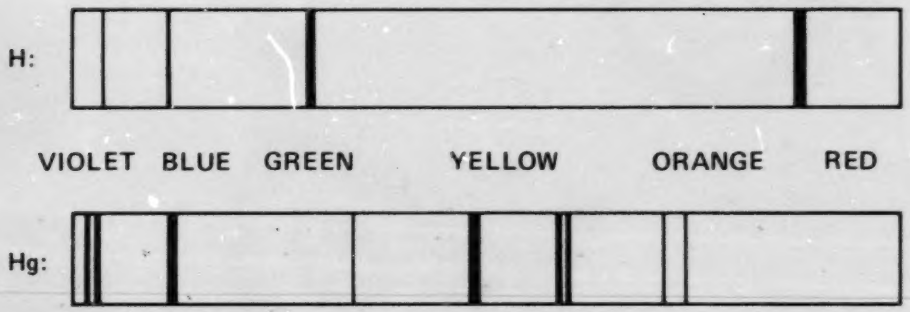


Figure 3 – Line spectra of hydrogen gas (top) and mercury vapor (bottom).

Emission-line spectra are easy to study in the laboratory. The brightest spectral lines of neon gas, for example, lie in the red region of the spectrum and thus account for the red color of neon signs. The spectrum of hot mercury vapor has prominent emission lines in the yellow, green, and blue regions of the spectrum (fig. 3), accounting for the greenish color of mercury-vapor street lamps. Sodium-vapor street lamps, on the other other hand, appear yellow because the brightest emission lines of hot sodium vapor lie in the yellow region of the spectrum. In emitting different colors of light to form an emission-line spectrum, a hot gas must be giving off energy. To keep the gas hot, an outside source of energy, such as an electric current, must be supplied. This energy is then radiated away in the form of light.

At the fairly high temperatures required to make gases radiate an emission-line spectrum, molecules are broken down into atoms. It is therefore, the atoms in the gas that are responsible for giving off the colors of light observed. Since different gases are composed of different kinds of atoms, it follows that different atoms produce different patterns of bright spectral lines when they receive energy from an outside source. The emission-line spectrum of a given hot, thin gas is, therefore, characteristic of the *atoms* making up that gas.

3. Absorption lines. If atoms can give off light when the necessary energy is supplied by an outside source, can they *absorb* light as well? The answer is "Yes," although demonstrating this effect in the laboratory is more difficult than demonstrating the emission of light. To show the effect, white light (containing all the colors of the spectrum) is passed through a cool gas and then through a slit into a prism. If the gas were not present in the path of the white light, a continuous spectrum would be observed on the other side of the prism. With the gas in the path of the white-light beam, however, the spectrum emerging from the prism is no longer continuous. Rather, *certain colors are missing* from the spectrum, as revealed by the presence of dark *absorption lines* crossing the continuous

spectrum of the white light. At the position of an absorption line, little or no light of that color is present in the spectrum. Obviously, atoms in the gas must be absorbing those colors of light that are missing in the spectrum. A pattern of dark absorption lines crossing an otherwise continuous spectrum is called an *absorption-line spectrum.*

We now come to a fundamental and very important point. The colors of light absorbed by a gas are the *same* colors emitted by the gas when it is supplied with an outside source of energy. Whether the atoms in the gas emit or absorb light depends upon the physical conditions in the gas. The emission of light requires either high temperature or an outside source of energy, whereas an absorption spectrum is observed only when a continuous spectrum of light is passed through the gas. In either case, however, the *pattern* of spectral lines is the same for a given gas – that is, for a given atom. Thus, an absorption-line spectrum, like an emission-line spectrum, can be used to identify the kind of atoms (and hence chemical elements) present in the gas being observed.

4. Spectroscopic analysis. What do we find if two different gases, corresponding to two different elements, are mixed together and the spectrum of the mixture is observed? If white light is passed through the mixture, the absorption spectrum will display dark absorption lines due to *both* of the gases (that is, *both* sets of absorption lines will be observed). Similarly, if the mixture is heated to a high temperature or supplied with an outside source of energy, emission lines characteristic of both component gases will appear in the emission-line spectrum of the mixture. It is clear that, if many different gases are mixed together, the spectrum will be very complicated since it contains the spectral lines of all the elements making up the mixture. An example of such a complicated spectrum is presented in figure 4, which shows the spectral lines of a hot mixture of the rare gases helium (He), neon (Ne), and argon (Ar), which are themselves elements. In this negative reproduction, the bright emission lines appear dark; therefore, figure 4 actually shows how the absorption spectrum of this mixture would look. (The numbers indicate the wavelengths of the spectral lines and may be ignored for the moment.)

However, it is possible to disentangle the various patterns of spectral lines in such a mixture. This is done by comparing the patterns with catalogs that list the colors of light absorbed and emitted by the atoms of all the chemical elements. Scientists can then recognize groups of spectral lines in the spectrum of a mixture as being the lines characteristic of previously studied elements. In this way, the composition of an unknown mixture of elements can be found from a study of its spectrum – a process known as *spectroscopic analysis.* For example, from a knowledge of the spectra of helium, neon, and argon, we could verify that the spectrum of figure 4 is, in fact, a mixture of these gases. Spectroscopic analysis is very important in astronomy since we have no means of carrying out laboratory experiments on the material of which stars are made. We must, therefore, rely on an analysis of starlight to determine the chemical composition of the stars.

C. The Light from the Stars

1. Basic facts about the stars. Stars are hot, glowing spheres of gas at great distances from the Sun, which itself is a typical star. Many of the brightest stars in the sky have Latin or Arabic names, such as Polaris (the North Star) and Aldebaran (the Eye of the Bull). Others

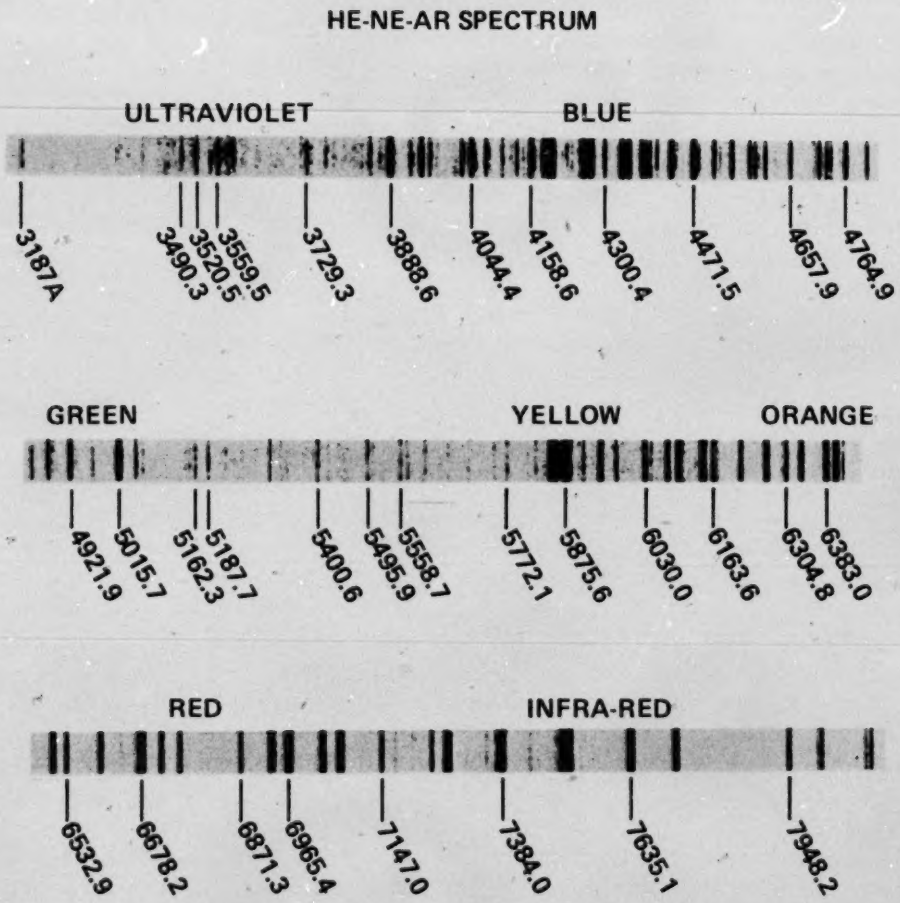


Figure 4 – Composite spectrum of the rare gases helium (He), neon (Ne), and argon (Ar). The numbers indicate the wavelengths of the spectral lines in Angstrom units.

are named by giving their brightness rank, in terms of a number or Greek letter, within the star group or constellation in which they appear – for example, β Canis Majoris or 31 Cygni. Still others are known only by their numbers in various catalogs, such as HD13147.

The temperature at the surface of a star can range from a few thousand degrees to perhaps as much as 100 000 degrees Kelvin, depending upon the kind of star. Stars are gaseous throughout, but the pressure and density of the stellar material, as well as the temperature, increase rapidly toward the center of the star. Temperatures at the centers of stars are many millions of degrees – so hot that nuclear reactions occur. The energy released by nuclear reactions is responsible for the entire energy output of the stars into space. The nuclear energy filters up through the interior layers of the star, eventually reaching the surface in the form of light. The starlight we can see and analyze comes mainly from the outermost layers of the star, near the surface.

The properties of the gas composing the star change gradually as we move from the surface of the star into the interior. For many purposes, however, we may picture a star as consisting of two distinct parts: a hot, opaque *sphere of highly compressed gas* and a cooler gaseous *atmosphere* by which the sphere is surrounded. Nearly all the light received from the star comes from its main body – the radiating sphere of highly compressed gas. Very little light is emitted from the somewhat cooler atmosphere. Instead, atoms in the stellar atmosphere absorb certain colors of light from the continuous spectrum of radiation emitted by the main body of the star, so that we observe an absorption spectrum.

2. The spectrum of the Sun. The first stellar spectrum to be studied was that of the Sun. Even simple observations of the sort performed by Newton reveal the presence of numerous faint *absorption lines* in the solar spectrum, and one may well wonder why Newton himself did not notice them. (Some believe this was because Newton had poor eyesight!)

The first detailed investigation of the solar specrum was carried out in the early 19th century by Fraunhofer. This scientist made an extensive catalog of the solar absorption lines and recorded their positions in the spectrum. Fraunhofer used capital letters of the alphabet to label the strongest lines, calling them the A-line, the B-line, etc. The cause of these absorption lines was unknown in Fraunhofer's day, but, in the middle of the 19th century, it was shown that absorption lines are created when a continuous spectrum of light is passed through a layer of gas. It was natural to assume that solar lines are formed in the same way and to ask whether the atoms responsible for the absorption of light in the solar atmosphere are the same as those on the Earth. Comparisons of the solar spectrum with the laboratory spectra of gases quickly showed that most of the solar absorption lines are due to elements known on the Earth. Figure 5 shows, for example, a portion of the yellow region of the solar spectrum; the two strong absorption lines in this region were called the D-lines by Fraunhofer. Laboratory experiments revealed that these two absorption lines are due to sodium vapor, showing that sodium atoms must be present in the solar atmosphere. Within a few years, several dozen terrestrial elements had been shown to be present on the Sun, although, even today, many hundreds of lines in the complicated solar spectrum remain unidentified.

It is interesting to note that 19th century scientists could not identify a particular set of absorption lines in the solar spectrum with those of any element known on Earth. It was

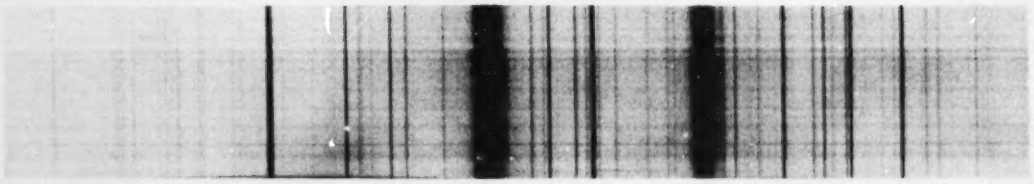


Figure 5 – A portion of the yellow region (5880 Å to 5900 Å) of the solar spectrum. The two strong absorption lines (Fraunhofer D-lines) are due to sodium vapor.

therefore proposed that these lines were caused by an element present on the Sun but not on the Earth. This new element was called "helium," from the Greek word "helios" meaning sun. Later, helium was found to exist on Earth, being present in significant quantities in deposits of natural gas. We now know that, next to hydrogen, helium is the most abundant element in the universe.

3. Stellar spectra. Study of the spectra of the brightest stars also began in the 19th century. Even from the earliest observations, carried out visually, it was clear that all stellar spectra do not look alike and that only a small fraction resemble that of the Sun. By the last decade of the 19th century, progress in the photography of faint objects permitted large numbers of stellar spectra to be photographed and classified. Figure 6 presents a representative sample of stellar spectra and shows how different they can be in general appearance. Photographs do not show the entire spectrum, but only the portion from the green (right-hand edge) to the violet (left-hand edge). Each spectrum is labeled by the name of the star in the right-hand margin. The letters and numerals at the left of the figure indicate the *spectral type* of the star, based on the overall appearance of its spectrum. Several spectral lines of hydrogen (H), belonging to the Balmer series, are identified. The spectrum of the Sun is not shown in figure 6, but it is roughly similar to that of the star β Com. Although emission lines can be seen in the spectra of some stars, the great majority of stellar spectral lines are absorption lines. These spectra generally confirm our view that certain colors of light are being absorbed from the continuous spectrum of the main body of the star by atoms in the cooler stellar atmosphere.

We now turn to some new questions: why is there such a wide range of types of stellar spectra, and why do most look so different from that of the Sun? For example, some of the spectra in figure 6 show strong lines of hydrogren, whereas others show these lines more weakly or hardly at all. Does this mean that stars differ widely in the amount of hydrogen they contain? Not necessarily; in fact, most stars contain much the same proportions of hydrogen and other elements.

To understand this point, let us return to a discussion of spectroscopic analysis. We have seen that different mixtures of elements have different spectra and that the presence of spectral lines due to a particular element is proof that the element is present in the mixture. There is, however, another important fact that we have not mentioned until now: the

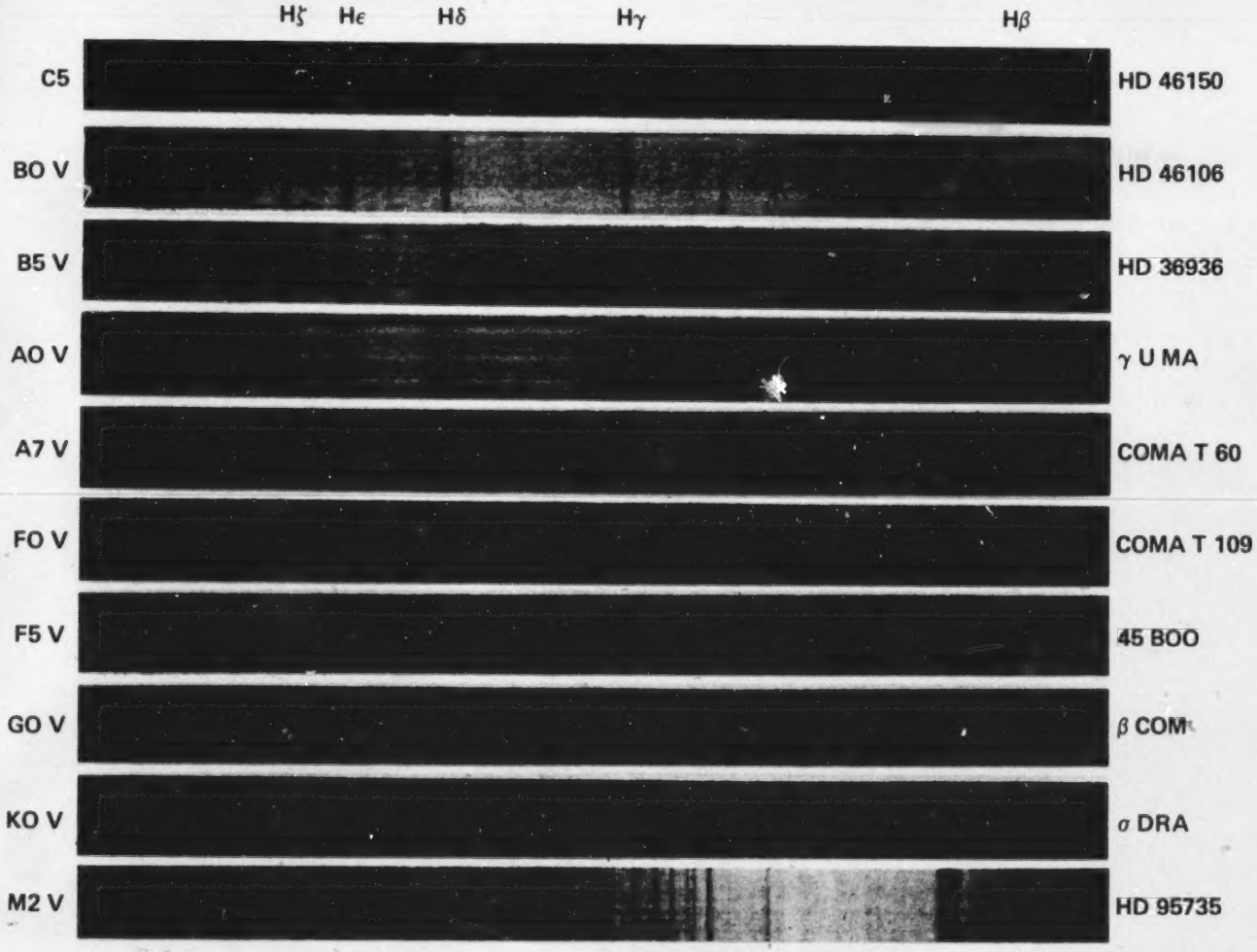


Figure 6 – The sequence of stellar spectral types.

spectrum of a gas composed of different elements depends not only on *what elements* are present, but also on the *temperature* of the gas. That is, the spectrum of a given mixture of elements can look different at different temperatures. The reasons for this *temperature effect* are discussed in later sections of this publication. Briefly, collisions between the atoms in the gas are more important at high temperatures than at low temperatures. The relative importance of collisions between the atoms – which may, for example, lead to ionization of the atoms – turns out to be a key factor in the appearance of the spectrum. Therefore, when the spectral lines of a given element are weak or missing in a spectrum, we must be cautious in our conclusions. This observation may mean that the element in question is present in low concentration, or not present at all, in the mixture of elements being observed. However, it may also mean that the temperature of the gas is not in the range of values that favors the absorption of those particular colors of light by the element in question.

In the late 19th century, many astronomers believed that the variety of stellar spectra was due to actual differences in composition among the stars. By the 1920s, however, advances in physics enabled scientists to take into account the temperature effect. The scientists were then able to predict how the spectrum of a given mixture of gases would look at different temperatures. They quickly found that almost all of the differences between various types of stellar spectra could be explained by the temperature differences of the stars. The proportions of elements making up the absorbing atmospheres of the stars proved to be roughly the same from star to star and very similar to the proportions of elements in the Sun's atmosphere. Most stellar spectra look different from that of the Sun, mainly because most stars are either hotter or cooler than the Sun. In figure 6, the stars whose spectra are shown have been arranged in order of decreasing temperature, with the hottest (perhaps 50 000 K) at the top and the coolest (about 3000 K) at the bottom. The Sun has a temperature of about 6000 K.

It is interesting to note that, in the first half of the 19th century, the French philosopher Auguste Comte speculated that human beings would never be able to understand the nature and composition of the stars. By the end of that century, spectroscopy had already begun to provide this information. In terms of the numbers of nuclei present, about 90 percent of a typical star is hydrogen. Helium makes up nearly all of the remaining 10 percent. The elements heavier than hydrogen and helium usually account for only 1 or 2 percent of the matter in the atmosphere of a star.

CHAPTER II

INTERMEDIATE AND ADVANCED TOPICS

A. More About Light: Electromagnetic Radiation

1. The electromagnetic spectrum. Evidence that light travels in *wave* dates from the 17th and 18th centuries. Examples of wave motion are given in basic physics classes: water waves, waves in a metal spring, sound waves in air, and others. (A wave sent down a jump rope, generated by shaking one end of the rope, is one of the simplest waves to produce and observe.) A wave is diagrammed schematically in figure 7. The time required for one complete up-and-down oscillation of the wave is the period P. The frequency f is closely related to the period; f is the number of cycles or oscillations the wave completes per second, so that $f = 1/P$. The speed v of the wave is the speed with which the "crests" of the wave move, and the wavelength λ (lambda) is the distance between two crests. One crest of the wave, traveling with speed v, covers a distance λ in the time $P = 1/f$; therefore,

$$\lambda f = v \qquad \text{(for a general wave)}$$

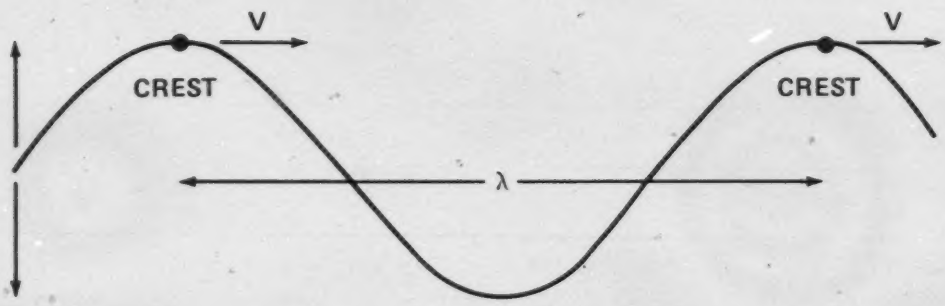


Figure 7 – Schematic diagram of a wave.

Early experiments showed that the wavelengths of visible light waves are extremely small. Therefore, it is convenient to define a special unit of length, called the *Angstrom unit*, in connection with light-wavelength measurements; one Angstrom = 1 Å = 1/100 000 000 cm = 10^{-8} cm. The *color* of visible light depends on its wavelength. Violet light, with a wavelength of about 4000 Å, is the shortest wavelength we can see. Green light has a wavelength of about 5000 Å, and the longest wavelengths visible are those of red light, about 7000 Å. To every color of light in the spectrum there corresponds a definite

wavelength, and vice versa. The range of wavelengths between 4000 Å and 7000 Å is the *visible spectrum*; we cannot see wavelengths shorter or longer than these.

We now know that light waves are just a special case of a more general kind of wave called an *electromagnetic wave*. Such waves are produced whenever electric charges are accelerated, and they are characterized by rapidly varying electric and magnetic fields of force. Electromagnetic waves of any wavelength can be generated, but only those with wavelengths in the range 4000 Å to 7000 Å can be detected by our eyes. All electromagnetic waves, regardless of wavelength, travel through empty space at the extremely high speed c = 300 000 km/s (where 1 km = 1000 m). This speed c is the "speed of electromagnetic radiation," or, more loosely, the "speed of light." The relation between the wavelength, frequency, and speed of an electromagnetic wave is thus

$$\lambda f = c \qquad \text{(for an electromagnetic wave)}$$

The range of all possible wavelengths of electromagnetic radiation is called the *electromagnetic spectrum* (fig. 8). Historically, waves in different parts of the spectrum were detected and investigated in some cases before it was realized that all are basically the same kind of radiation. It is significant that the visible spectrum of light waves occupies only a small portion of the electromagnetic spectrum. To the short-wavelength side of the visible spectrum lie the ultraviolet, X-ray, and gamma-ray regions; to the long-wavelength side lie the infrared, microwave, and radio regions. Different means of detection must be used for waves in different regions of the spectrum. Many astronomical bodies, such as stars, radiate energy over a wide range of wavelengths and not just in the visible part of the spectrum. Therefore, all wavelengths of the electromagnetic spectrum must be analyzed to give us a true picture of the physical conditions in astronomical bodies.

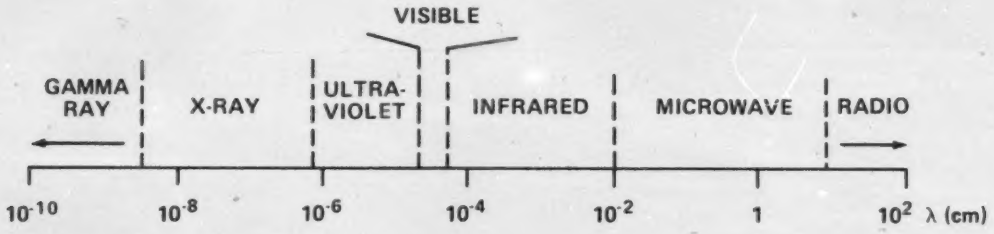


Figure 8. – The electromagnetic spectrum.

2. The quantum theory of electromagnetic radiation. One of the most striking discoveries of 20th century physics is the *dual nature* of electromagnetic radiation: under some conditions, electromagnetic radiation behaves like a wave, whereas, under other conditions, it behaves like a stream of particles called *photons*. The tiny amount of energy carried by a photon is called a *quantum* of energy (plural – *quanta*), from the Latin word for "little bundle." The quantum theory of radiation was first introduced in 1900 by the great

German physicist Max Planck to explain certain properties of radiating bodies. Planck showed that the quantum of energy E carried by a photon of radiation was proportional to the frequency f of that radiation:

$$E = hf \qquad \text{(the Planck formula)}$$

where h is a fundamental constant of physics known as the *Planck constant*. (If E is measured in ergs and f in Hertz, or cycles per second, then $h = 6.624 \times 10^{-27}$ erg-s.) From our previous equation relating the wavelength, speed, and frequency of an electromagnetic wave, it is clear, for example, a wave of very short wavelength will have a very high frequency. Thus, a gamma-ray or X-ray photon will carry a quantum of energy that is much greater than that of a visible-light or microwave photon because the frequency is much higher.

The quantum theory of radiation is of great importance in atomic physics. We have seen that atoms emit and absorb only certain colors – that is, wavelengths – of electromagnetic radiation. To each wavelength, or frequency, there corresponds a photon carrying a quantum of energy whose size is given by the Planck formula. Thus, we can also say that atoms emit or absorb *quanta* of electromagnetic energy in the form of photons. This quantum description draws attention to the fact that, in emitting or absorbing photons, the atom is losing or gaining energy in certain definite amounts.

3. Kirchhoff's laws. We have previously seen that there are basically three types of spectra: continuous, emission-line, and absorption-line. A more complete statement of the conditions under which these types of spectra are observed was given by the German physicist Kirchhoff in 1859:

- *First law:* A hot, glowing, solid body, liquid, or highly compressed gas emits a continuous spectrum of electromagnetic radiation.
- *Second law:* A hot, glowing, thin gas emits a bright-line spectrum of electromagnetic radiation.
- *Third law:* If a continuous spectrum of electromagnetic radiation is passed through a layer of this gas, the gas will absorb certain wavelengths of radiation from the continuous spectrum, and an absorption spectrum will be observed; the wavelengths absorbed are the same as the ones emitted when the gas is heated.

We have seen that atoms emit or absorb only certain characteristic wavelengths of radiation. However, in a solid, liquid, or highly compressed gas, the atoms are jammed close together and interfere with each other's natural radiating processes. The result is that, under the conditions of the first law, the emitted radiation is actually a blend of spectral lines – in effect, a continuous spectrum. To observe the effects of individual atomic emission or absorption, we must experiment with a thin gas. In a thin gas, the atoms are relatively far apart and do not substantially interfere with each other during the emission or absorption of radiation. If such a thin gas is heated, we observe an emission-line spectrum characteristic of the atoms of that gas (second law). If continuous radiation is passed through the gas, we observe an absorption-line spectrum characteristic of individual atoms (third law); in this

case, however, the layer of gas must be sufficiently cool that its own emitted radiation does not cancel the absorption effect to be observed.

B. Hydrogen: the Simplest of All Atoms

1. The Balmer lines. As might be expected, hydrogen has the simplest spectrum of any atom. The first spectral lines of hydrogen to be studied lie mainly in the visible region of the spectrum, although numbers of them crowd together toward a limiting wavelength in the near ultraviolet (fig. 9). This series of lines is called the *Balmer* series, after the Swiss schoolteacher who first found the formula for the spacing of the lines. The first few lines are labeled by Greek letters. Hα, the first line in the Balmer series, lies in the red at 6562 Å. The second Balmer line, Hβ, lies in the green at 4861 Å; Hγ lies in the blue at 4340 Å; and so on. Although the first few Balmer lines lie in the visible spectrum, the remainder fall in the ultraviolet, crowding closer and closer together until finally they cannot be recognized as individual lines. There are no Balmer lines with a wavelength less than 3646 Å; this limiting wavelength is the Balmer *series limit*.

Noting that the lines appeared to have a very regular spacing, Balmer found – by trial and error in 1885 – a formula that predicts their wavelengths. If the first, second, third, etc., Balmer lines are numbered by integers $n = 3$, $n = 4$, $n = 5$, etc., as shown in figure 9, then the Balmer formula for the wavelength λ_n of the nth Balmer line is

$$\frac{1}{\lambda_n} = R\left[\frac{1}{4} - \frac{1}{n^2}\right], n = 3, 4, 5, \ldots,$$

where R is a fundamental constant of physics called the *Rydberg constant*: $1/R = 911.4$ Å.

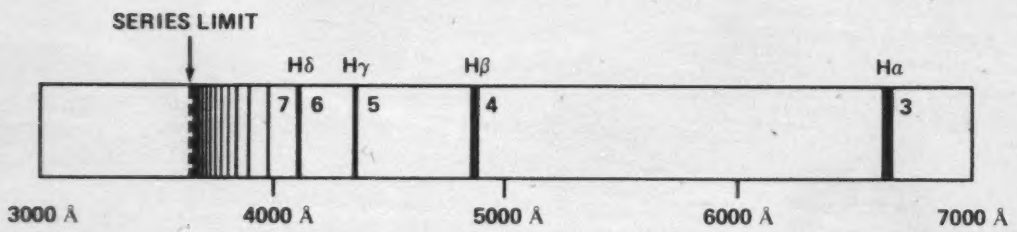


Figure 9 – The Balmer lines of hydrogen, which crowd together toward a series limit of 3646 Å in the ultraviolet.

2. The Bohr electron orbits. Until 1913, no one understood why the Balmer formula should be correct. In that year, the famous Danish physicist Niels Bohr showed that the electron in the hydrogen atom cannot move in any orbit around the nucleus, but only in certain circular orbits numbered 1, 2, 3, and so on (fig. 10). The electron cannot emit or absorb radiation as long as it moves in one of these *Bohr orbits*. Each Bohr orbit

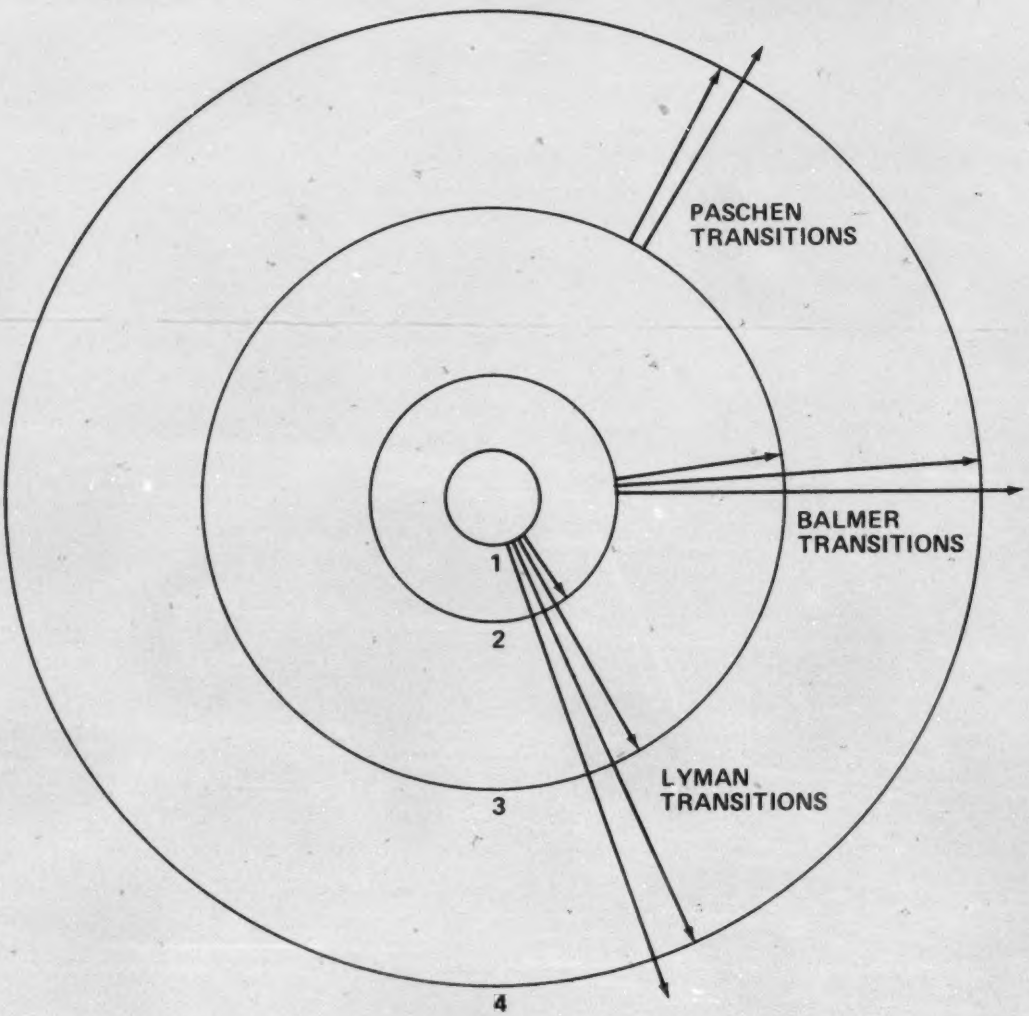


Figure 10 – Bohr electron orbits in hydrogen. Transitions (electron jumps) from the 1, 2, and 3 orbits up to larger orbits give rise to the Lyman, Balmer, and Paschen series of absorption lines; transitions in the reverse directions correspond to emission lines.

corresponds to a definite *energy state* of the hydrogen atom. When the electron is in the $n = 1$ orbit, the atom is in its lowest state of energy. Other orbits represent higher energy states; the higher the n value and the larger the orbit, the higher the energy state of the atom.

Thus, the hydrogen atom can emit or absorb energy only if the electron jumps between different Bohr orbits. If an electron jumps from a large, high-energy orbit down to a smaller, low-energy one, the atom *emits* electromagnetic radiation; Bohr found that the amount of energy carried away by this radiation exactly equals the *difference* in energy of the two orbits involved. Using the quantum description, we may also say that the atom emits a photon whose energy is equal to the difference in energy of the two orbits. This quantum of energy corresponds, by the Planck formula, to an electromagnetic wave of a definite frequency and wavelength; therefore, the downward jump of the electron is revealed by the presence of an emission line in the spectrum. The hydrogen atom can also *absorb* quanta of energy in electron jumps from smaller to larger orbits. However, the electron cannot jump from a smaller to a larger orbit unless the energy of the incident radiation exactly matches the energy difference between two orbits. That is, a photon of just the right energy is required to generate an electron jump up to a particular larger orbit. Therefore, radiation of most frequencies will not lead to any absorption electron jumps. But, if a continuous spectrum of radiation is passed through a gas of hydrogen atoms, there will be some radiation at just the right frequencies to cause various upward jumps of the electrons. The atoms will then absorb these frequencies or wavelengths from the beam of continuous radiation, and we will observe an absorption spectrum. Since the wavelengths absorbed or emitted are determined by the energy differences between the various Bohr orbits, this theory explains the fact that the emission and absorption spectra of hydrogen gas have the same patterns of spectral lines.

Bohr further showed that the radius of the nth Bohr orbit is proportional to n^2 and that the energy associated with this orbit is proportional to $1/n^2$. Using the Planck formula to relate the energy difference of the orbits n and m to the wavelength λ_{mn} of the radiation absorbed or emitted in electron jumps between them, Bohr was able to prove the *Rydberg formula*

$$\frac{1}{\lambda_{mn}} = R\left[\frac{1}{m^2} - \frac{1}{n^2}\right] \qquad \text{(the Rydberg formula)}$$

where it is assumed that m is less than n. This relationship had been established earlier by Rydberg from observations of several series of hydrogen spectral lines. Electron jumps beginning from or ending on the $m = 1$ orbit and where $n = 2, 3, 4, \ldots$, give rise to the *Lyman series* of spectral lines in the ultraviolet region of the spectrum. Electron jumps beginning from or ending on the $m = 2$ orbit, with $n = 3, 4, 5, \ldots$, give rise to the *Balmer series* of spectral lines (see figs. 6 and 9). Other series of lines correspond to electron jumps beginning from or ending on still larger orbits – for example, the Paschen series ($m = 3$) in the infrared (see fig. 10 for the case of absorption transitions).

Historically, the *Balmer* lines of hydrogen have been the most important because only this series of lines has members lying in the visible region of the spectrum. Let us further

examine the Balmer absorption transitions, for which the electron must initially be in the 2 orbit. The atom can then absorb radiation in an electron jump from the 2 orbit to the 3 orbit, the wavelength of this radiation being that of the first Balmer line, Hα. The second Balmer line, Hβ, corresponds to an electron jump from the 2 orbit to the 4 orbit, and so on. The integers numbering the Balmer lines in figure 9 give the number of the higher orbit to which the electron is jumping when the atom absorbs radiation of a wavelength corresponding to that spectral line. These wavelengths are given by the $m = 2$ case of the Rydberg formula, which then reduces to the Balmer formula.

If we let n "go to infinity" in the Balmer formula – that is, if we drop the term involving $1/n^2$ – we can calculate the wavelength of radiation required to remove the electron to an infinitely large orbit and so *ionize* the atom. A simple calculation gives this wavelength as $4 \times (1/R) = 3646$ Å approximately; this is the *series limit* wavelength for the Balmer series. Balmer wavelengths of radiation longer than this can *excite* an electron from the 2 orbit up to a higher energy orbit but cannot ionize the atom. Any wavelength of radiation shorter than the Balmer series limit, however, can ionize a hydrogen atom with its electron in the 2 orbit; in this case, the freed electron is left with some residual kinetic energy.

3. Excitation and ionization of hydrogen by collisions. We have seen that radiation can *excite* hydrogen atoms by moving their electrons up to higher energy orbits and can also ionize them. It is important to recognize, however, that *collisions* between the atoms (and with electrons of other atoms) can have the same effect.

In a gas of hydrogen atoms, energy is constantly being exchanged among the atoms by collisions. The kinetic energy of the atoms as a whole represent part of this energy, but some energy is also available to excite the electrons in the atoms up to higher energy orbits. At low temperatures, there is little collisional energy available for excitation, and nearly all of the atoms have their electrons in the lowest energy, $n = 1$ Bohr orbit. As the temperature is increased, however, the proportion of atoms with electrons in the $n = 2, 3, 4$, and higher orbits is increased as well. Above about 10 000 K, so many of the electrons are in highly excited orbits that *ionization by collision* becomes important since electrons far from the nucleus are easily removed by collisions with other atoms or electrons. At very high temperatures, nearly all of the gas is ionized; there are then no emission nor absorption lines visible in the spectrum as there are hardly any netural atoms left to do the absorbing and emitting!

C. Stellar Spectral Types

1. Classification of stellar spectra. The spectra shown previously in figure 6 are representative of the many types encountered; that is, most stellar spectra look rather similar to one of these representative types. The early work on stellar spectra consisted of crude attempts to put the spectra into various classes, or *types*, on the basis of their general appearance: the number of spectral lines present, the positions of the lines, and the strengths of the lines. These types were called A, B, C, and so on, after the letters of the alphabet.

By the early 1900s, when many photographs of spectra were available for study, two facts emerged. First, it was found that earlier workers had put into separate classes many spectra that were actually similar so that some classes could be eliminated or included in others. Second, and more important, it was found that the remaining types could be arranged in a row, or sequence, in which the various spectral features varied smoothly from one end of the sequence to the other. Some spectra do not fit into this one-dimensional classification scheme, but perhaps 90 percent of all spectra do in fact fit in somewhere in the so-called *sequence of spectral types*, or *spectral sequence*. The letters of the spectral types run in this order: O-B-A-F-G-K-M. (The standard mnemonic for the spectral sequence is "*O*h, *B*e *A F*ine *G*irl! *K*iss *M*e!") The O-type spectra are characterized by weak Balmer lines of hydrogen and the lines of ionized helium, as well as a few lines of other highly ionized elements. The B-type spectra display lines of neutral helium, and the Balmer lines are stronger than in type O. The A-type spectra have the strongest Balmer lines of any spectral type. In F-type spectra, the Balmer lines are weaker than in type A and the spectrum begins to grow more complex, with many lines of both neutral and ionized metals starting to appear. The Balmer lines fade steadily through the spectral types G, K, and M, as lines due to neutral metals and, finally, molecules dominate the spectrum. Within a given spectral type, the spectra are further graded in appearance on a scale of zero to nine. For example, an A5 spectrum is midway in appearance between an A0 spectrum and and F0 spectrum. The Sun has a type G2 spectrum. Types O, B, and A are customarily called "early" spectral types and types K and M "late" spectral types, as it used to be thought (incorrectly) that the spectral sequence represented successive stages of stellar evolution.

2. Strength of the Balmer absorption lines: a clue to temperature. The fact that nearly all stellar spectra can be arranged in a one-dimensional sequence strongly suggests that most stars differ in only one respect. In the 1920s, it was shown that differences in *temperature* can account for nearly all differences between the spectra in the spectral sequence of figure 6. These temperature differences lead to different degrees of *excitation* and *ionization* of the atoms in the stellar atmosphere and, therefore, to pronounced variations in the appearance of the spectra. The strength of the hydrogen Balmer lines, in the visible part of the spectrum, provides a good illustration of the effects of excitation and ionization. (The "strength" of an absorption line is proportional to the *percentage* of continuous radiation absorbed at the wavelength in question; the actual *amount* of continuous radiation present at that wavelength does not matter.)

Consider first the cases of the O stars, which are now known to have surface temperatures of 25 000 K or higher. At such high temperatures, because of the violent collisions between the atoms, hydrogen is mostly *ionized* into electrons and protons. In the atmospheres of the hot O stars, then, there are very few neutral hydrogen atoms left to absorb radiation at wavelengths corresponding to the Balmer or any other series of spectral lines. The Balmer lines are weak in O-star spectra, therefore, because hydrogen is largely ionized in the atmospheres of these stars. Consider next the B stars, which are somewhat cooler – in the range of 11 000 K to 25 000 K. At these lower temperatures, a smaller fraction of the hydrogen atoms is ionized by collisions. Thus, the Balmer absorption lines are stronger in B stars than in O stars.

In stars of still lower temperature, a new effect becomes important. It is still true that the lower the temperature, the smaller the fraction of hydrogen atoms that are ionized. But, at lower temperatures, collisions between the atoms are less energetic and thus also less effective in *exciting* the electrons up to higher energy orbits. The result is that most of the atoms have their electrons in the $n = 1$ orbit of lowest energy. Such atoms can absorb *Lyman* wavelengths of radiation, which correspond to transition from the $n = 1$ orbit up to higher orbits (fig. 10), but not the Balmer wavelengths. Therefore, as we examine spectral types F (about 7000 K), G (about 5000 K),, type K (about 4000 K), and M (about 3000 K), the Balmer lines fade steadily; fewer and fewer atoms are "candidates" for Balmer absorption. As it happens, the competition between the effects of excitation and ionization yields the most favorable conditions for Balmer absorption at temperatures typical of the A-type stars. If astronomers of the early 1900s could only have studied the absorption lines of hydrogen in other regions of the electromagnetic spectrum (unfortunately blocked out by the Earth's atmosphere), they would have realized at once that the range of strengths of the Balmer lines in stellar spectra is not due primarily to actual hydrogen abundance differences among stars. Rather, this variation in Balmer line strengths is due largely to the combined effects of excitation and ionization of the hydrogen atoms in the atmospheres of the stars – effects which depend strongly upon temperature.

3. Complex atoms and ions in stellar spectra. An atom or ion with two or more electrons is said to be *complex*. Because the motion of each electron about the nucleus is influenced by the presence of all the other electrons, a complex atom or ion cannot be described in terms of a collection of simple, Bohr-type orbits similar to those used to describe the hydrogen atom. Rather, the structure of a complex atom is critically dependent upon the number of electrons actually present. The most striking aspect of this dependence is the tendency of the electrons to occupy distinct groupings called *shells* at various average distances from the nucleus. The "shell structure" of a complex atom is responsible for its chemical properties and is discussed more fully in chemistry textbooks.

Like hydrogen, however, a complex atom or ion has certain definite *states of energy* corresponding to different arrangements of the electrons around the nucleus. Electromagnetic radiation is emitted or absorbed in *transition* from one atomic energy state to another. In the case of complex atoms, unfortunately, there is no simple Rydberg-type formula for the wavelengths of radiation emitted or absorbed. These must be determined from laboratory measurements or from difficult, approximate calculations based on the quantum theory of atomic structure.

Complex atoms are also like hydrogen in that they can be excited to states of higher energy by either radiation or collision. However, when a complex atom is ionized by the removal of an electron, the resulting system – consisting of the nucleus together with the remaining electrons – is a positive *ion* which itself is capable of emitting and absorbing electromagnetic radiation. It is important to note that the *spectrum of an atom is entirely different from the spectrum of any of its ions*, which also *differ completely* from each other. The general appearance of the spectrum depends mainly upon the number of electrons still attached to the nucleus.

Because collisions between the atoms are usually the main cause of ionization, the degree of ionization of a complex atom in a stellar atmosphere is an important indicator of *temperature*. If the lines of a neutral atom are missing in a stellar spectrum, it does not mean that there is none of that element in the stellar atmosphere. It may be that the atoms are ionized one or more times. In a complete spectral analysis, scientists look for both the spectral lines of a neutral atom and the spectral lines of its various ions. Since the nuclei are unaffected by the ionization process, the true abundance of an element is determined by the *number of nuclei* of that element present in the atmosphere. The number of electrons still attached to those nuclei – that is, the stage of ionization – depends upon the temperature (and also the density) of the gas. Unfortunately, the spectral lines of most positive ions fall in the ultraviolet or even X-ray regions of the spectrum because large amounts of electromagnetic energy are involved in the electronic transitions, and such wavelengths are blocked out by the Earth's atmosphere. A complete analysis of stellar spectra thus demands observations made above the Earth's atmosphere. As a result of the United States space program, and particularly the efforts of the National Aeronautics and Space Administration, many satellite observations of ultraviolet stellar spectra have already been made. A vigorous program of such observations will greatly increase our knowledge of stellar properties in the future.

It is gratifying that analyses of the spectral lines of complex atoms and ions lead to very much the same conclusions about stellar compositions and temperatures as an analysis of the Balmer lines of hydrogen. For example, the neutral atoms of most common metals have outer electrons that are rather easily excited to higher energy orbits so that they are easily ionized as well. The atoms of magnesium, calcium, iron, and other elements exist as positive ions in the atmospheres of the hottest stars; the spectral lines of these metal ions generally fall in the ultraviolet and do not, therefore, appear on photographs of the visible region of the spectrum. Only by spectral type F is the temperature sufficiently low to allow these atoms to exist in the neutral state. In late-type stars, the lines of neutral metal atoms are very numerous in the visible region of the spectrum and dominate its appearance. Recall that, by contrast, the Balmer lines of hydrogen are very weak in the coolest stars because of insufficient excitation up to the $n = 2$ level, even though hydrogen is far more abundant than all other elements combined.

GLOSSARY

absorption– the removal of colors of light (or wavelengths of electromagnetic radiation) from a beam of light or from the radiation of an atom or a substance.

Ångström unit (symbol: Å)– a small unit of length, equal to 1/100 000 000 or 10^{-8} cm, used to measure wavelengths of light waves.

atmosphere, stellar– the layer of absorbing gas surrounding the compressed dense central portion of a star.

atom– the smallest unit of a chemical element, composed of a heavy nucleus surrounded by orbiting electrons.

Balmer lines– the wavelengths of radiation absorbed or emitted by the hydrogen atom in the visible region of the spectrum.

Bohr orbits– the circular orbits of possible electron motion in the hydrogen atom.

complex– as applied to an atom or ion, containing two or more electrons.

compound, chemical– a substance composed of two or more chemical elements.

continuous– as applied to a spectrum, one with no missing colors of light or wavelengths of electromagnetic radiation.

electron– a light, negatively charged particle; in an atom, electrons orbit around the atomic nucleus.

element, chemical– one of the substances (92 of which occur naturally) that cannot be broken down into other substances by chemical reaction; composed of atoms.

emission– the giving off of colors of light (or wavelengths of electromagnetic radiation) by an atom or substance.

excitation– the process by which atomic electrons are raised from lower to higher energy levels through collisions, or the absorption of electromagnetic radiation.

hydrogen– the lightest and most abundant of all elements; hydrogen atoms contain only one electron.

ion– a positively charged atomic system created by the removal of one or more electrons from an atom.

ionization– the removal of electrons from an atom by collisions or electromagnetic radiation.

Kelvin temperature scale– measures temperature above absolute zero; water freezes at 273 K and boils at 373 K.

Kirchhoff's laws– laws describing the physical conditions under which continuous, emission-line, or absorption-line spectra are observed.

line, spectral– a line of color, or a wavelength of electromagnetic radiation, appearing in (emission) or missing from (absorption) a spectrum.

molecule– the smallest unit of a chemical compound, composed of two or more atoms.

neutron– a heavy nuclear particle carrying no electric charge.

nucleus– the small, dense central part of an atom, composed of protons and neutrons, around which atomic electrons orbit.

photon– a "particle" of electromagnetic radiation carrying a quantum unit of energy proportional to the frequency of the radiation.

proton– a heavy, positively charged nuclear particle.

quantum (plural: quanta)– a small, but definite, amount of energy carried by a photon of electromagnetic radiation.

sequence, spectral– the arrangement of stellar spectral types in order of decreasing temperature: O-B-A-F-G-K-M.

spectroscopy– the analysis of light or electromagnetic radiation in terms of its component colors or wavelengths.

spectrum (plural: spectra)– the array of colors of light or wavelengths of electromagnetic radiation present in a source of light or radiation.

spectrum, electromagnetic– the array of all possible wavelengths of electromagnetic radiation, from gamma rays (shortest wavelength) to radio waves (longest wavelength).

spectrum, visible– the range of wavelengths of electromagnetic radiation, from 4000 Å to 7000 Å, visible to the human eye.

type, spectral– the classification of a stellar spectrum, based on its overall appearance, into one of the categories O-B-A-F-G-K-M of the spectral sequence.

wave, electromagnetic– a wave generated by the acceleration of electric charges and characterized by rapidly varying electric and magnetic force fields; moves with the speed c = 300 000 km/s in a vacuum.

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(D) *Exploration of the Universe*, by Abell, George; Holt, Rinehart, and Winston, 1969 (2nd edition). One of the many excellent astronomy texts at the college-freshman level, with a good discussion of stellar spectra.

(E) *Catalog of Photographs and Slides from the Hale Observatories*, obtainable from the California Institute of Technology Bookstore, Pasadena, California. Lists many commercially available slides of solar and stellar spectra.

CLASS EXERCISES

(I) Basic Topics

(a) Spectrum identification laboratory activity: equip students with (1) diffraction gratings or student spectroscopes, and (2) strips of black construction paper together with colored chalk; the teacher will need (3) gas-tube light sources, such as neon, mercury, sodium, etc., and (4) a chart in the colors of the emission spectra of the gases in the light sources. With lights off, have students sketch on their black strips with the colored chalk the emission lines they see from the glowing gas tubes. When several spectra have been sketched, ask students to identify the gases by comparison with the spectra illustrated in the laboratory color chart. If individual diffraction gratings are available, students may take these home and observe mercury- and sodium-vapor street lights through them.

(b) Solar absorption lines: you will again need gratings or spectroscopes. After masking a suitably placed classroom window with opaque paper (and drawing the shades on the others), allow a beam of sunlight to enter the room and fall upon a white wall or white sheet of paper. Try to observe with gratings or spectroscope the solar absorption lines in the sunlight reflected from the white surface. (The lines may be faint and hard to see, but, to avoid eye damage, be sure to caution students never to observe the Sun directly, even with a diffraction grating or spectroscope.)

(II) Intermediate and Advanced Topics

(a) Wave motion: Use a ripple tank, a long metal spring, or other means to demonstrate various forms of wave motion. Make sure students see and understand what is meant by the speed, frequency, and wavelength of the waves, and show that the equation relating these quantities holds for all of these types of wave motion.

(b) The electromagnetic spectrum: Assign outside reading or projects on the ways in which gamma rays, X-rays, microwaves, and radio waves are detected. Discussion of this reading can be related to the use of X-rays in medicine, the cooking of food by microwaves, and the production and reception of radio waves by antennas. In the same vein, you might ask students why they would be unable to get a suntan behind a glass window, even though they would feel the heat of the Sun. (Answer: glass absorbs ultraviolet wavelengths of radiation, which produce a suntan, but transmits the heating infrared wavelengths.)

(c) Wave nature of light: If equipment is available and students have sufficient background, demonstrate the wave nature of light by single- and double-slit interference and diffraction experiments. (Suitable slits may be made by ruling fine scratches in graphite-blackened microscope slides.)

QUESTIONS AND ANSWERS

(I) Basic Topics

A. What are Atoms?

Q: What is the difference between a molecule and an atom? (A: An atom is the smallest unit of a chemical element. A molecule, the smallest unit of a chemical compound, consists of atoms joined together.)

Q: What is the structure of an atom? (A: A small, dense, central nucleus, surrounded by a cloud of orbiting electrons.)

Q: What particles make up the atomic nucleus? (A: Protons and neutrons.)

Q: What holds an atom together? (A: A force of electrical attraction between the positively charged nucleus and the negatively charged electrons.)

Q: Which requires more energy – to break apart atoms into electrons and nuclei or to break apart nuclei into protons and neutrons? (A: Much more energy is required to break apart nuclei.)

B. Light and Spectroscopy

Q: Who first showed that white light is a mixture of all the colors of the rainbow, and how did he do it? (A: Isaac Newton, by spreading out the light with a prism which separates the colors.)

Q: Why is it convenient to use a slit source of light in spectroscopy experiments? (A: Because the wide band of colors is easy to study.)

Q: What kinds of light sources produce a continuous spectrum? (A: Hot, glowing, solid bodies or highly compressed gases.)

Q: How is an absorption-line spectrum produced? (A: By passing white light, containing a continuous spectrum of colors, through a gas.)

Q: Are the colors of light absorbed by an atom of a gas the same as, or different from, the colors emitted by that atom? (A: The same.)

Q: How can the composition of a sample of gas be found by spectroscopic analysis? (A: By comparing the spectrum of the gas, in either emission or absorption, with laboratory spectra of the elements.)

C. The Light from the Stars

Q: What kind of spectral lines do we see in the spectra of most stars – emission lines or absorption lines? (A: Absorption lines.)

Q: Do stellar spectra look mostly the same? (A: No; there are very great differences between them.)

Q: What is the main cause of the differences between the various types of stellar spectra? (A: Differences in the temperature of the stars.)

Q: What conclusions have astronomers drawn about the chemical composition of the stars? (A: The chemical composition of most stars is roughly the same as that of the Sun: about 90 percent hydrogen, 10 percent helium, and several percent of the remaining elements, by nuclear count.)

(II) Intermediate and Advanced Topics

A. More About Light: Electromagnetic Radiation

Q: How are the wavelength, frequency, and speed of a wave related? (A: The wavelength times the frequency equals the speed of the wave.)

Q: What is an Angstrom unit, and why is it a convenient unit in which to measure wavelengths of light? (A: One Angstrom equals 1/100 000 000 centimeter, convenient because light wavelengths are so small.)

Q: What is the "visible spectrum?" (A: Those wavelengths of radiation in the range 4000-7000 Å that the human eye can see.)

Q: What are the names for the various regions of the electromagnetic spectrum? (A: In order of increasing wavelength: gamma-ray, X-ray, ultraviolet, visible, infrared, microwave, and radio.)

Q: How is the energy of a photon of electromagnetic radiation found from the frequency of the radiation? (A: The energy equals the frequency multiplied by Planck's constant h.)

Q: Kirchhoff's laws describe the conditions under which the three basic kinds of spectra are observed; what are these kinds of spectra? (A: Continuous, emission-line, and absorption-line.)

B. Hydrogen: The Simplest of All Atoms

Q: What series of spectral lines of hydrogen lie in the visible region of the spectrum? (A: The Balmer series.)

Q: What quantities does the Balmer formula allow us to calculate? (A: The wavelengths of the Balmer lines of hydrogen.)

Q: What are the "Bohr orbits?" (A: The circular orbits of permitted electron motion in the hydrogen atom.)

Q: How does the electron in a hydrogen atom jump when the atom emits or absorbs radiation? (A: In emission, the electron jumps down from a larger, higher energy orbit

to a smaller, lower energy one; in absorption, the electron jumps from a smaller to a larger orbit.)

Q: What is the number of the orbit from which, or to which, the electron jumps if the hydrogen atom absorbs or emits Balmer wavelengths of radiation? (A: The $n = 2$ orbit.)

C. Stellar Spectral Types

Q: What is the order of spectral types in the spectral sequence? (A: The order is O-B-A-F-G-K-M.)

Q: How does the strength of the Balmer lines of hydrogen vary with spectral type? (A: Weak in type O, stronger in type B, strongest in type A, decreasing in strength through types F, G, K, and M.)

Q: What explains the weakness of the Balmer lines in the hottest stars (type O) and the coolest stars (type M)? (A: In type O stars, the hydrogen is mostly ionized; in type M stars, few hydrogen atoms have their electrons excited up to the $n = 2$ orbit by collisions.)

Q: If we cannot detect the spectral lines of a neutral atom in a stellar spectrum, does that necessarily mean that there is little or none of that element in the stellar atmosphere? (A: No, not necessarily; the atoms may be ionized one or more times.)

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